

05-08-06

Chapter 12/13/14 starts here!

Solutions & their Behaviors

Mixture - two or more substances physically combined. -> everything in it keeps its own properties in which they retain their individual props.

• 3 types: 1) suspension - a mixture which has large particles -> can see w/ the naked eye. -> particles settle.

ex. rocks & water.

2) colloid - a mixture that has an intermediate particle size -> can NOT be seen with a naked eye & they don't settle out. [can see under a microscope]

ex. mayo, milk, etc.

3) solution - a mixture that has the smallest particle size -> can NOT be seen with the naked eye & they don't settle out.

to tell difference...
Tyndall Effect

• Tyndall effect - shine light on mixture -> if can see light: colloid - particles are big enough to shine light.

-> if can't see the light: solution

... looks can be deceiving!

• 2 classifications: 1) homogeneous (uniform) } throughout mixture.
2) heterogeneous (not uniform)

Solutions

• 2 parts: 1) solute - gets dissolved

2) solvent - does dissolving -> for this class: usually water -> making aqueous sol.

Nature of solutes

ex $\xrightarrow{H_2O}$ (dissociate into sep. ions)
- ionic $\rightarrow Cu(NO_3)_2 (s) \xrightarrow{H_2O} Cu^{2+} + 2NO_3^-$
- molecular (covalent): ex $CuH_2O \xrightarrow{H_2O} CuH_2O (aq)$

★ • when dissolved in H_2O , ionic dissociate into their respective ions.
break a part

• no dissociation in molecular solutes -> just become aqueous.

electrolytes - any compound that dissolves in H_2O & allows for electricity to be conducted.

05-09-06

class demonstration

• bulb -> water: faint light
-> sucrose (aq): less faint than water
-> calcium chloride (aq): Bright light.

an electrolyte - any substance that dissolved in H_2O & allows the sol. to conduct electricity.

non electrolyte (ie. sucrose) - any substance that dissolves in H_2O & allows the sol. to conduct elec.

• $CaCl_2 \rightarrow Ca^{2+} + 2Cl^-$ (dissociates b/c it's an ionic bond)
=> allow H_2O to conduct electricity

• $C_{12}H_{22}O_{11} \rightarrow C_{12}H_{22}O_{11}$ -> doesn't dissociate.

- no change - just a carbohydrate
=> current isn't carried through.

• water -> polar molecule -> reason why ions in ionic compound dissociate

• sucrose dissolves in water -> b/c of hydrogen bonding => sugar becomes polar

• only polar ~~molecules~~ ^{solutes} will dissolve in polar solvents. [like dissolves in like]

05-09-06

(Dissolving)

Solution Process:

1) Break solute-solute interactions (requires E)

2) Break solvent-solvent interactions (also requires E)

3) Form solvent-solute interactions (releases E)

→ endothermic

overall
endothermic
process
[usually]

→ depends on
how much E is
released

* solvation (if dealing w/ water → hydration) - when solute particles are completely surrounded by solvent particles in a soln.

Rate of Solution Process

1) Stir it up ⇒ more KE, more collisions b/w solute & solvents (Kinetic Molecular Theory)

⇒ brings fresh solvent w/ solute.

(ex. Dunkin' Donuts 3-sugar coffee)

2) Heat it up ⇒ adds KE (b/c temp. is raised) → more collisions w/ solute.

3) Increased surface area ⇒ more SA → more contact w/ solute & solvent.

Solubility - the amount of solute that dissolves in a specific amount of solvent at a specific temp.

⇒ temperature dependent (more things will dissolve @ a higher temp.)

* When solvent holds a max. amount of solute → saturated solution (ie sponge)

* When a solvent holds less than max. amount of solute → unsaturated solution

* When a solvent holds more than the max. amount of solute at a given temperature → supersaturated solution

05-10-06

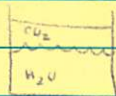
Factors affecting solubility

1) Temperature: as temp ↑, sol ↑ (for the most part) } for solids

* for gases: as temp ↑, sol ↓

2) Polarity: "like dissolves in like" ⇒ polar solutes dissolve in polar solvents. (why ionic compounds dissolve in water)
⇒ nonpolar solutes dissolve in nonpolar solvents.

3) Pressure (ONLY for gases): as p ↑, sol ↑ = more gas collisions with liquid → more gas dissolved



→ Henry's Law

* miscible - any two substances [mainly liquids] can be dissolved in each other in any proportion

ex. alcohol & water

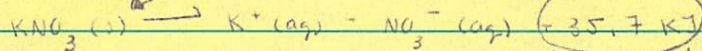
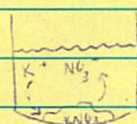
* immiscible - two substances [mainly liquids] can not be dissolved in each other.

ex. oil & water.

05-12-016

Equilibrium in saturated solutions

[sorry!]



endothermic

heat of sol. (the ΔH change associated with a solute dissolving)

$\Delta H_{\text{soln.}}$

→ diss. occ. each solution.

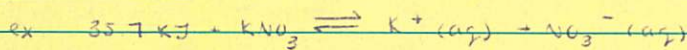
$$\Delta H_{\text{soln.}} \begin{cases} = (+) \text{ [endothermic process]} \\ = (-) \text{ [exothermic process]} \end{cases}$$

$$\begin{aligned} \text{ex. } & \text{KNO}_3 = +35.7 \frac{\text{kJ}}{\text{mol}} \\ & \text{CO}_2 = -19.91 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

• in a saturated sol. ^{an} equilibrium (Eq) → a dynamic state (always moving!) in which two opposing processes (rxns) [ie dissolving & precipitating] happen at equal rates.

- b/c molecules const. to move → always solute reaching & combining w/ solvent.

Le Chatelier's Principle: If a system at equilibrium is stressed, the Eq. will shift to relieve that stress.



change in temp, pressure, concentration, etc.

• If temp. increases → Eq. shifts right → more dissolves (wants to relieve stress!)

• If more K^+ ions are added → Eq. shifts left → more precipitate forms & T is increased. $[\text{NO}_3^-]$ will decrease → helps to relieve stress

→ can connect to solubility curve.

(gases → are exothermic - most solids are endothermic)

Concentrations ([])

$$\% \text{ by mass} = \frac{\text{mass of solute}}{\text{total mass of soln}} \cdot 100\% = \frac{\text{mass of solute}}{\text{mass of solute} + \text{solvent}} \cdot 100\%$$

05-15-06

Concentrations (I)

• % by mass = $\frac{\text{mass solute}}{\text{mass of sol} + \text{mass of solvent}} \cdot 100\%$

what is

Ex: what is [%] of 9.80 g NaCl dissolved in 210 mL of H_2O ?

mass solute \rightarrow 9.80 g NaCl

mass solution \rightarrow can't add mL & g...

• $D_{H_2O} = 1.0 \text{ g/mL}$

$\Rightarrow \frac{9.80 \text{ g NaCl}}{9.80 \text{ g} + 210 \text{ g}} \cdot 100\%$

= 4.45% NaCl

\rightarrow b/c water is a liquid \rightarrow can't use $PV = nRT$

\rightarrow in college \rightarrow will want initials & date of when sol. was made.

most important of the 3 • Molarity (M) - ex. 1.0 M $CuSO_4$ \Rightarrow 1.0 mol of $CuSO_4$ in 1 L of soln.

$\rightarrow \frac{\text{moles of solute}}{\text{L of soln.}}$

\rightarrow also seen as $\left(\frac{\text{moles}}{\text{L}}\right)$.

Ex: what is the [M] of 7.00 g of Na_2SO_4 in 500 mL of soln?

\rightarrow never given a vol. of H_2O .

- convert g to mol

$\rightarrow 7.00 \text{ g } Na_2SO_4 \cdot \frac{1 \text{ mol } Na_2SO_4}{142.1 \text{ g } Na_2SO_4} = 0.049 \text{ mol } Na_2SO_4$

$\rightarrow 500 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}} \rightarrow 0.5 \text{ L of soln}$

$\Rightarrow \frac{0.049 \text{ mol } Na_2SO_4}{0.5 \text{ L}} = \text{0.0980 M } Na_2SO_4$

if can NOT convert b/w M & m

• M = mass

• M = molar mass

• Molality (m) = $\frac{\text{moles of solute}}{\text{kg of solvent}}$

\rightarrow important in calculating the Δ in freezing pt & BP.
 \rightarrow used in bond constant.

Ex: what is the [M] of 9.80 g of NaCl in 210 mL H_2O ?

$9.80 \text{ g NaCl} \cdot \frac{1 \text{ NaCl}}{58.5 \text{ g NaCl}} = 0.167 \text{ mol NaCl}$

$210 \text{ mL} \cdot \frac{1 \text{ g}}{1 \text{ mL}} = 210 \text{ g} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 0.210 \text{ kg } H_2O$

$\Rightarrow \frac{0.167 \text{ mol NaCl}}{0.210 \text{ kg } H_2O} = \text{0.797 M NaCl}$

if can convert b/w [%] & [M]

05-23-06

Vapor Pressure - the pressure exerted by a vapor of a liquid

- all liquids have a vapor pressure.

Equilibrium of Vapor Pressure [Eg VP]

Ex. $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g})$ [not equal amounts, but equal rates of condensation & evaporation]
 $\Rightarrow \text{E} + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g})$

• as temp \uparrow , VP \uparrow

Boiling - changing from liquid to gas phase within & at the surface of a liquid.

Boiling Point - the temperature at which vapor pressure of a liquid and atmospheric pressure are equal.

• high VP \Rightarrow evaporates faster.

• volatile (liquids) - evaporate readily (high VP)

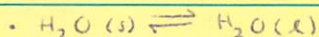
• nonvolatile (liquids) - evaporate slowly (low VP)

• by lower P, can make liquids boil at lower temperature [In class demo, w/ boiling H_2O at room temperature]

05-25-06

Freezing - changing from liquid state to solid state.

Freezing Point (FP) - the temperature at which solid & liquid phases of a substance are in equilibrium.



Phase Diagram - compares the three states of a substance at any temperature & pressure.

• 3 general sections \rightarrow (S) solid, (L) liquid, & (G) gas.

• Sublimation - solid \rightarrow gas

• Deposition - gas \rightarrow solid.

• Critical temperature - the temperature at which only vapor phase exists for a substance @ any P.

• Triple point - T & P at which all 3 states of a substance are in equilibrium.

05-26-06

Colligative property - [like a physical property] it depends on the amount of solute particles
 - not the nature of the solute

- Vapor Pressure (VP) - VP goes down as [solute] goes up
 → b/c the solute particles are solvated (when solvent molecules are completely surrounded by solute particles)
 - these attractive forces keep the solvent from forming vapor

- Boiling Point (BP) - BP goes up as [solute] goes up
 → b/c w/ solute added, VP ↓ ⇒ more E is needed to make a solvent's VP equal to atmospheric pressure

• Molality = $\frac{\text{moles solute}}{\text{kg solvent}}$ → won't change with temperature.

• vol. of solution changes w/ temperature ⇒ molality Δ w/ temperature

• $\Delta T_b = K_b m$
 $\underbrace{\Delta T_b}_{\text{change in boiling temp.}} = \underbrace{K_b}_{\text{Molal BP constant}} m$
 $\underbrace{K_b}_{\text{solvent specific}}$

ex. pure H₂O = 100.0°C

1 M = 100.51°C

2 M = 101.02°C

$$K_b (\text{for H}_2\text{O}) = 0.51 \frac{^\circ\text{C}}{m} *$$

- Freezing Point (FP) - as [solute] increases, FP decreases

→ b/c solute particles are blocking the solvent molecules from forming their highly organized solid structure ⇒ they need to move slower
 to get a solid to form, so temperature FP goes down

• $\Delta T_f = K_f m$
 $\underbrace{\Delta T_f}_{\text{change in freezing temp.}} = \underbrace{K_f}_{\text{Molal FP constant}} m$

ex. pure H₂O = 0.0°C

1 M = -1.86°C

2 M = -3.72°C

$$\Rightarrow K_f (\text{for H}_2\text{O}) = -1.86 \frac{^\circ\text{C}}{m} *$$

05-30-06

Ex. What is the new FP & BP of a soln. of 30.0g sucrose in 2.14 kg of H_2O ?

$$30.0g \text{ } C_{12}H_{22}O_{11} \cdot \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342g \text{ } C_{12}H_{22}O_{11}} = 0.0877 \text{ mol } C_{12}H_{22}O_{11}$$

$$M = \frac{0.0877 \text{ mol } C_{12}H_{22}O_{11}}{2.14 \text{ kg } H_2O} = 0.041 M \text{ } C_{12}H_{22}O_{11}$$

Δ in FP

$$\Delta T_f = K_f M \Rightarrow (-1.86^\circ C/M)(0.041 M \text{ } C_{12}H_{22}O_{11}) = \boxed{-0.0762^\circ C}$$

$$\Delta T_b = K_b M = (0.51^\circ C/M)(0.041 M \text{ } C_{12}H_{22}O_{11}) = \boxed{0.0209^\circ C}$$

Δ in BP

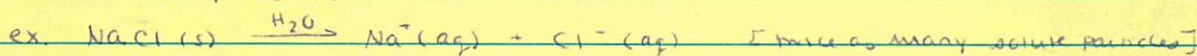
$$\Rightarrow \text{new BP} : 100^\circ C + 0.0209^\circ C$$

$$= \boxed{100.021^\circ C}$$

* sucrose
nonvolatile
nonelectrolyte.

Electrolytes + ΔT_b + ΔT_f

\rightarrow ionic compounds dissociate in water.



\Rightarrow more in contact w/ H_2O molecules

• BP \rightarrow increase

• FP \rightarrow decrease.

• for ionic compounds...

$$\Delta T_b = K_b M(n) \quad \cdot (n) = \# \text{ of solute particles}$$

$$\Delta T_f = K_f M(n)$$

